

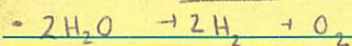
Chpt 9 HWK

P 277 Q 3a
287 Q 3, 4
294 Q 2, 4
296 Q 314, 15, 17
297 Q 26, 29

03-09-06

Chpt 9 starts here!

Chpt 9: Stoichiometry - mass relationships in a chemical rxn.



2 mol. of ^{water}~~hydrogen~~ 1 mol. of oxygen \rightarrow 2 moles. of hydrogen.

If wanted 2 moles. of $\text{O}_2 \Rightarrow$ 4 moles. of $\text{H}_2\text{O} \Rightarrow$ 4 moles. of H_2 .

If had 6.622×10^{23} moles. of $\text{O}_2 \Rightarrow 1.2 \times 10^{24} \text{H}_2 \Rightarrow 1.2 \times 10^{24}$ moles. H_2O .

$\underbrace{\hspace{10em}}_{\text{a mole of } \text{O}_2} \quad \underbrace{\hspace{10em}}_{2 \text{ moles of } \text{H}_2} \quad \underbrace{\hspace{10em}}_{2 \text{ moles of } \text{H}_2\text{O}}$

* knowing reaction \rightarrow can apply amount of moles.
 \Rightarrow molar ratios \rightarrow conversion factors

Mole to Mole Ex.

If 3 moles of O_2 are produced, how many moles of H_2O are needed?

$$3 \text{ mol } \text{O}_2 \cdot \frac{2 \text{ mol } \text{H}_2\text{O}}{1 \text{ mol } \text{O}_2} = 6 \text{ mol of } \text{H}_2\text{O}$$

molar ratio

- conversion factor comparing moles of reactants & products from the equation.

• easier to quantify in grams \rightarrow

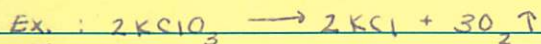
ex. $4.5 \text{ mol } \text{O}_2 \cdot \frac{16(?) \text{ gram } \text{O}_2}{1 \text{ mol } \text{O}_2} = \boxed{144 \text{ gm. } \text{O}_2}$

- easy to do vice versa \rightarrow mass to moles problems.

mass to
moles
problem \rightarrow

03-13-06

Mole-Mole:



• If 3.04 g of KCl are produced, how many moles of O_2 are produced?

$$3.04 \text{ mol KCl} \cdot \frac{3 \text{ mol O}_2}{2 \text{ mol KCl}} = 4.56 \text{ mol O}_2$$

molar ratio

Mole-Mass:

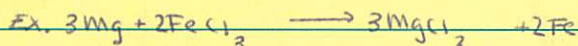
Ex: If 3.04 g KCl are produced, how many grams of O_2 are produced?

[cont. of ex. above]

$$4.56 \text{ mol O}_2 \cdot \frac{32 \text{ g O}_2}{1 \text{ mol O}_2} = 147.6 \text{ g O}_2$$

• $\text{mole A} \rightarrow \text{mole B} \rightarrow \text{mass of B}$
molar ratio molar mass ratio

Mass-Mole:



• If 2.06 g of Mg react with excess of FeCl_3 , how many ^{mole} ~~grams~~ of iron are produced?
 ↑
 so much extra, don't need to worry about it.

$$2.06 \text{ g Mg} \cdot \frac{1 \text{ mol Mg}}{24.3 \text{ g Mg}} \cdot \frac{2 \text{ mol Fe}}{3 \text{ mol Mg}} = 0.0565 \text{ mol Fe}$$

• If one doesn't have an excess ...

Ex: 2.06 g Mg & 3.50 g \rightarrow find amount of moles of Fe.

\rightarrow which # to use? = have to do certain equations [later in chpt]

Mass-Mass:

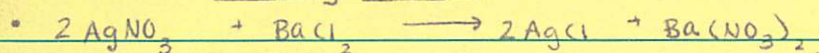
Ex: [cont. of ex. above] How many grams of Fe?

$$0.0565 \text{ mol Fe} \cdot \frac{55.8 \text{ g Fe}}{1 \text{ mol Fe}} = 3.15 \text{ g Fe}$$



• $\text{mass A} \rightarrow \text{mole A} \rightarrow \text{mole B} \rightarrow \text{mass B}$
molar mass molar ratio molar mass
 $(n) \text{ g A}$ $\frac{\# \text{ mole B}}{\# \text{ mole A}}$ $(n) \text{ g B}$

03-16-06 Limiting Reaction

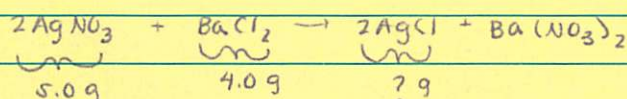


If 5.0 g of AgNO_3 react with an excess of BaCl_2 , how many g of AgCl are produced?

plenty of the
reactant

→ don't have to worry about it.

→ new question: If 5.0 g of AgNO_3 react with 4.0 g of BaCl_2 , how many g of AgCl are produced?



• limiting reactant (AKA limiting reagent) - the reactant that runs out 1st
→ thus stopping the rxn.

• excess reactant - the reactant in abundance that will not run out 1st.

Steps for a Limiting Reactant Problem

1) convert both amounts to moles.

Ex: • $5.0 \text{ g AgNO}_3 \cdot \frac{1 \text{ mol AgNO}_3}{170 \text{ g AgNO}_3} = 0.0294 \text{ mol AgNO}_3$ have *

• $4.0 \text{ g BaCl}_2 \cdot \frac{1 \text{ mol BaCl}_2}{208 \text{ g BaCl}_2} = 0.01923 \text{ mol BaCl}_2$ have *

2) Pick one reactant & use a molar ratio to see how much of the other reactant needs one.

Ex: $0.0294 \text{ mol AgNO}_3 \cdot \frac{1 \text{ mol BaCl}_2}{2 \text{ mol AgNO}_3} = 0.0147 \text{ mol BaCl}_2$ need

⇒ have more than that

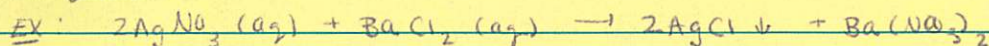
⇒ AgNO_3 is the limiting Reactant!

3) compare moles of reactant [need to what one has]

- if one has plenty

03-20-06

Limiting Rxn



(A) 5.0 g AgNO_3

(B) 4.0 g BaCl_2

produce how many grams of AgCl .

Step 1: convert both to moles.

• 5.0 g $\text{AgNO}_3 \rightarrow 0.0294 \text{ mol } \text{AgNO}_3$ (have)

4.0 g $\text{BaCl}_2 \rightarrow 0.0192 \text{ mol } \text{BaCl}_2$ (have)

Step 2: pick one reactant (A) & see how many moles of other reactant (B) are needed.

• $0.0294 \text{ mol } \text{AgNO}_3 \cdot \frac{1 \text{ mol } \text{BaCl}_2}{2 \text{ mol } \text{AgNO}_3} = 0.0147 \text{ mol } \text{BaCl}_2$ (needed)

Step 3: compare moles needed (B) to moles have (B) of reaction.

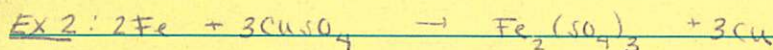
- if you have more "B" than you need \Rightarrow B is the excess & A is the limiting reactant.

for EX. \Rightarrow (A) limiting reactant & (B) is excess reactant.

- if you have less "B" than you need \Rightarrow B is the limiting reactant & A is excess.

Step 4: do all calculations with limiting reactant.

$0.0294 \text{ mol } \text{AgNO}_3 \cdot \frac{1 \text{ mol } \text{AgCl}}{1 \text{ mol } \text{AgNO}_3} \cdot \frac{143.5 \text{ g } \text{AgCl}}{1 \text{ mol } \text{AgCl}} = 4.22 \text{ g } \text{AgCl}$



• If 4.52 g of Fe react with 2.78 g CuSO_4 , how many grams of Cu are produced? (a)

• $4.52 \text{ g } \text{Fe} \cdot \frac{1 \text{ mol } \text{Fe}}{55.8 \text{ g } \text{Fe}} = 0.0807 \text{ mol } \text{Fe}$ have (b)

• $2.78 \text{ g } \text{CuSO}_4 \cdot \frac{1 \text{ mol } \text{CuSO}_4}{159 \text{ g } \text{CuSO}_4} = 0.0175 \text{ mol } \text{CuSO}_4$ have

- how much excess reagent is left over? (in grams)

• If pick Fe $\rightarrow 0.0807 \text{ mol } \text{Fe} \cdot \frac{3 \text{ CuSO}_4}{2 \text{ mol } \text{Fe}} = 0.12105 \text{ mol } \text{CuSO}_4$ need \Rightarrow don't have it!

$\Rightarrow \text{CuSO}_4$ is the LR.

• $0.0175 \text{ mol } \text{CuSO}_4 \cdot \frac{3 \text{ mol } \text{Cu}}{3 \text{ mol } \text{CuSO}_4} \cdot \frac{63.5 \text{ g } \text{Cu}}{1 \text{ mol } \text{Cu}} = 1.12 \text{ g } \text{Cu}$ (a)

(b) How much ER is left over?

$0.0807 \text{ mol } \text{Fe}$ have - [how much needed]

$\rightarrow 0.0175 \text{ mol } \text{CuSO}_4 \cdot \frac{2 \text{ mol } \text{Fe}}{3 \text{ mol } \text{CuSO}_4} = 0.0117 \text{ mol } \text{Fe}$ used

$\Rightarrow 0.0807 \text{ mol } \text{Fe} - 0.0117 \text{ mol } \text{Fe} = 0.0690 \text{ mol } \text{Fe} \cdot \frac{55.8 \text{ g } \text{Fe}}{1 \text{ mol } \text{Fe}} = 3.87 \text{ g } \text{Fe}$

03-23-06

$$\text{Percent Yield} = \frac{\text{actual yield (amount of product from experimentation)}}{\text{theoretical yield (calculated yield)}} \cdot 100\%$$

Ex: If 4.0 g of sodium sulfate reacts with an excess of calcium chloride, what is the percent yield of the ppt if only 0.800 g of it are produced?



$$4 \text{ g Na}_2\text{SO}_4 \cdot \frac{1 \text{ mol Na}_2\text{SO}_4}{142 \text{ g Na}_2\text{SO}_4} \cdot \frac{1 \text{ mol CaSO}_4}{1 \text{ mol Na}_2\text{SO}_4} \cdot \frac{136 \text{ g CaSO}_4}{1 \text{ mol CaSO}_4} = 3.84 \text{ g CaSO}_4$$

↳ theoretical yield.

$$\text{percent yield} = \frac{0.800 \text{ g}}{3.84 \text{ g}} \cdot 100 = 20.8\% \rightarrow 21\%$$